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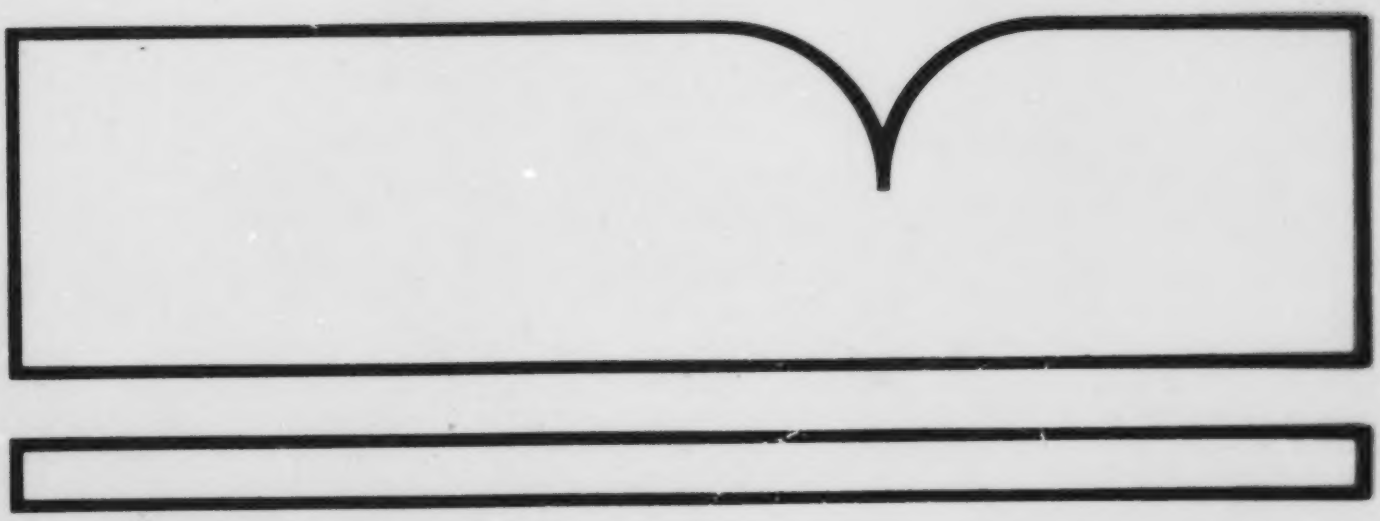
Movement of Phosphorus in Soil

Robert S. Kerr Environmental Research Lab.
Ada, OK

1984



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THE MOVEMENT OF PHOSPHORUS IN SOIL

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U.S. ENVIRONMENTAL PROTECTION AGENCY
ADA, OK 74820

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ABSTRACT

The disappearance of phosphate from soil solution results from a fast initial sorption reaction followed by a much slower transformation of phosphate to amorphous and crystalline precipitates of iron, aluminum, and calcium. The kinetics of the transformation have been determined at a phenomenological level and often shown to follow Elovich or first-order equations. Kinetic models are applied to laboratory data to describe step response of phosphorus and quasi-steady state response of phosphorus in soil columns. Results indicate a first-order kinetic equation gives an adequate description of phosphorus movement in soil columns when the water velocity is known.

Index words: miscible displacement, chemical kinetics, adsorption, precipitation

INTRODUCTION

Eutrophication of lakes and streams is a natural process which can be accelerated by the movement of nutrient rich water into waterways. Phosphorus (P) has been identified as the nutrient most likely limiting primary productivity in lakes and streams. A prerequisite to understanding the impact of land application wastewater treatment or the environmental impact of P fertilization is an assessment of the interaction of P and soil constituents.

The soil is a dynamic chemical and biological system in a state of constant flux. To describe the complex system, the paper begins with a description of the chemical forms of P and associated soil constituents and then the kinetics of the reaction. After describing the soil P interaction, P transport is compared to measured P activity in dynamic systems.

There are many forms of P which may be applied to the soil ranging from complex organics to simpler orthophosphates. Mineralization of many of the P compounds appear to be reasonably rapid (Blanchar and Hossner, 1969). The discussion presented here will be limited to the reaction and movement of orthophosphate.

CHEMICAL STABILITY MODELS

Most inorganic phosphates found in the soil can be classified into three groups: (1) those containing calcium phosphates; (2) those containing iron and aluminum phosphates; and (3) those combining with silicate materials. The relative importance of these compounds can be roughly correlated to pH of the soil environment. In acid soils, iron and aluminum phosphates are more common. In basic soils, calcium phosphates predominate.

By studying the solubility products of some of the more important iron, aluminum, and calcium compounds commonly found in soil, it is possible to quantitatively project stability diagrams for the equilibrium concentration of phosphorus in soil solution.

The equilibrium isotherms, plotted in Figure 1 for selected P compounds, were developed from Gibbs free energies in Table 1. The values selected for the Gibbs free energies were from single source (Sadiq and Lindsay, 1979) so that they would be internally consistent with each other. There is a wide range in the values reported in the literature, and the projected concentrations could range several orders of magnitude if the reported extremes were used. For this reason, using data that is internally consistent is important. Several assumptions were made in calculating Figure 1: (1) calcium was assumed to be 0.01 molar; (2) the ionic strength was assumed to be 0.026 mole liter^{-1} (ionic strength assumed equal to 0.013 times the electrical conductivity in millimhos cm^{-1}); (3) variscite ($AlPO_4 \cdot 2H_2O$) was assumed to be in equilibrium with amorphous aluminum hydroxide ($Al(OH)_3$); (4) strengite ($FePO_4 \cdot 2H_2O$) was assumed to be in equilibrium with iron hydroxide ($Fe(OH)_3$); and (5) in the pH range of interest (HPO_4^{2-}) and ($H_2PO_4^-$) were the only phosphate species significantly affecting the P activity. When possible, it would be better to know the activity of the iron, aluminum, calcium, and

oxidation state rather than estimate activity from an assumed associated compound. The stability model can then be used to estimate the equilibrium concentration of P in soil solution when pH and activities of Ca^{2+}, Fe^{2+}, Fe^{3+}, and Al^{3+} are known.

SORPTION-DESORPTION MODELS

There is confusion as to the terminology used in the literature describing the loss of phosphorus from solution. Some choose to describe their data using solubility product theory, as above, (van Riemsdijk et al. 1975; Bennett and Adams, 1976; Subbarao and Ellis, 1977; Lindsay, 1979; and Hsu, 1979) assuming a precipitation type of process. Other researchers (Holford et al. 1974; Shayan and Dovey, 1978; and Sibbesen, 1981) attempt to describe the process using sorption isotherms. The most commonly used sorption expression for phosphorus is the Langmuir equation (Langmuir, 1918) which can be written in its simplest form as

$$\frac{C}{S} = \frac{1}{bS_m} + \frac{C}{S_m} \qquad (1)$$

where:

C = solution concentration of phosphorus (mg/l)

S = adsorbed phosphate at concentration C (µg P/g soil)

S_m = adsorption of maximum which is equivalent to a monolayer of phosphate on the adsorbing surface (µg P/g soil)

b = a constant

The second most common equation is the Freundlich equation

$$S = K C^{n} \qquad (2)$$

where:

S = phosphate adsorbed (µg P/g soil)

K = constant

C = equilibrium concentration of phosphorus (mg P/l)

n = constant

The Freundlich equation has been shown to be equivalent to a multiple layer Langmuir equation. Both of the methods have successfully described the disappearance of P in soil after a given time for equilbration (Berkheiser et al. 1980). Both of these equations, as generally written, assume phosphorus reactions are instantaneous and the soil is at equilibrium. If these equations alone were used to predict the movement of phosphorus in soil, one would expect phosphorus movement to appear as if it were moving in a chromatographic column. The chemical, when applied, would move through the soil as a slug. This type of description does not appear consistent with most reported column data in literature or batch sorption data which shows the influence of equilibrating time.

Consider sorption data of Stuanes (personal communication 1981) (Figure 2). The data presented is after twenty-four hours of equilibration. The soil suspension was centrifuged and the equilibrating solution changed. The amount of sorption was calculated by measuring the reduction in P in the supernatant liquid. The process was repeated ten times to evaluate both sorption and desorption. The data shows a nearly irreversible reaction. Equations 1 and 2 are generally written as if the system remains in local equilibrium. One can hypothesize, from the data in Figure 2, that either sorption isotherms, as they are generally written, are not adequate for describing P sorption or one must also include the kinetics of the reaction and assume the experimental data of Stuanes does not represent an equilibrium condition to have a useable model.

Stuanes' data does not conflict with equilibrium solubility models since the mass of sorption or desorption will be different dependent on initial conditions. However, Figure 3 shows the timed response to several applications of P. The system depicted in Figure 3 does not appear to be at equilibrium. This indicates the solubility product approach would also need to include a kinetic term to adequately describe P movement.

KINETIC MODELS

The disappearance of P from soil solution occurs as a fast initial reaction followed by a much slower reaction. Investigators believe sorption is the primary process during the initial rapid reaction, and precipitation to relatively insoluble phosphate controls the slow reaction (Sawhney, 1977). The quick response takes place in the first few hours followed by a gradual change for a long period of time.

Methods of describing this response varies widely among researchers and there appears to be little agreement as to the appropriate approach. One group will describe the response assuming sorption processes control the reaction; another group believes the process is controlled by the stability of the chemicals and the rate of nucleation or crystal growth. A third group attempts to describe the process empirically. Most researchers who use the sorption process base their description on the assumption that at least a portion of the sorption sites are not at equilibrium. de Camargo et al. 1979 describe their assumption as two types of sites: type I sites are assumed to be in equilibrium following the equation

$$\frac{\partial S_1}{\partial t} = \frac{\theta}{\rho} K_1 C^N - K_2 S_1 \qquad (3)$$

type II sites are assumed to not be at equilibrium following the equation

$$\frac{\partial S_2}{\partial t} = \frac{\theta}{\rho} K_3 C^N - K_4 S_2 \qquad (4)$$

with S_1 and S_2 the amount of P adsorbed at the respective sites. K_1 and K_2 are the forward and backward rate coefficients for type I sites and K_3 and K_4 the forward and backward rate coefficients for type II sites. To be

at equilibrium, $\frac{\partial S_1}{\partial t}$ must equal zero. Equation 3 then becomes the same as a Freundlich equation $S_1 = KC^N$. The retardation of the pollutant as pointed out by de Camargo et al. (1979) will be $R = 1 + (\rho\ N\ K\ C^{N-1}/\theta)$. Overman et al. (1978) used a similar approach but assumed the reaction followed a linear equilibrium condition rather than a Freundlich equation.

Several (Novak et al. 1975; Mansell et al. 1977; and Enfield et al. 1975) assumed equilibrium could be described by either the Langmuir equation or the Freundlich equation, but the system had not yet reached equilibrium. Novak et al. 1975 assumed the mass transfer is proportional to a concentration driving force (C - C*) where C* is the equilibrium condition described by the Langmuir equation; the mass transfer rate is then described using first-order kinetics as

$$\frac{\partial S}{\partial t} = K\ (C - C^*) \qquad (6)$$

Mansell et al. 1977, on the other hand, used equilibrium described by a Freundlich equation. Enfield et al. 1976 assumed the process was diffusion limited following the equation

$$S_{avg} = F(C) - \frac{6\ F(C)}{\pi^2} \sum_{j=1}^{n} \frac{1}{j^2}\ \exp\ (-\ Kj^2\ \pi^2\ t/r^2) \qquad (7)$$

where:

F(C) = Langmuir or Freundlich equilibrium isotherm

S_{avg} = average concentration of P in the soil particle

K = diffusion coefficient

r = radius of assumed spherical soil particle where phosphorus is being sorbed

Several of the approaches proposed by researchers (Novak et al. 1975; Sawhney, 1977; Mansell et al. 1977a; Mansell et al. 1977b; Overman et al. 1978; Fiskell et al. 1979; Enfield et al. 1975; and Enfield et al. 1976) may be adequate to describe the kinetics of the sorption process but do not appear to be adequate to describe the desorption process as presented in Figure 2.

Enfield et al. 1981 took the approach that sorption was an instantaneous reversible reaction following the Langmuir equation but the dominant factor controlling the reaction kinetics was the rate of nucleation and crystal growth considering solubility product theory. Their equation describing the kinetic reaction in soil was

$$S = \frac{S_m\, b\, C}{1 + b\, C} + \sum_{i=1}^{n} \int_o^t a_i\, (C - CE_i)\, dt \qquad (8)$$

where:

a_i = rate coefficients

CE_i = equilibrium concentration for a specific compound

i is an index for the numerous compounds which might form

As the equation is written, several compounds could be forming simultaneously each following first-order kinetics. Formation of multiple compounds at a given time has been identified recently by Freeman and Rowell, 1981. As the equilibrating solution goes down, the number of compounds forming would be reduced with an associated disociation of the more soluble compounds. The equation would project the concentration should fall in steps; for example, going from DCPD-OCP-TCP-HA where the solution would appear to stay near the DCPD line if the sum of the rates of formation of OCP, TCP, and HA were less than the rate of dissolution of DCPD. This type of response has been reported by Subbarao and Ellis 1977.

TRANSPORT MODELS

The general equation describing the process of simultaneous diffusion, mass flow, and sorption of a reactive solute in a soil in one dimension is

$$\frac{\partial C}{\partial t} = D \frac{\partial^2 C}{\partial X^2} - V \frac{\partial C}{\partial X} - \frac{\rho}{\theta} \frac{\partial S}{\partial t} \qquad (9)$$

where C is the concentration of P in the soil solution ($mg^{-1}l$), t is time (hrs), D is a coefficient combining diffusion and hydrodynamic dispersion ($cm^2 hr^{-1}$), V is the interstitial velocity ($cm\ hr^{-1}$) which is the flux of solution divided by the volumetric water content θ ($cm^3\ cm^{-3}$), ρ is the soil bulk density ($g\ cm^{-3}$), S is the amount of P sorbed (refers to the total amount adsorbed and/or precipitated) per unit mass of soil ($\mu g\ g^{-1}$), and X is the coordinate along the flow path (cm). There have been numerous solutions to this equation for a variety of boundary conditions and rate terms. Three typical solutions will be examined here. Assumptions for the first solution are:

(1) θ, ρ, D and temperature remain constant during the experiment;

(2) the sorption-desorption processes remain in local equilibrium following the linear relationship $S_1 = k_d C$;

(3) an irreversible precipitation process takes place only when it has previously been sorbed on the solid phase following first-order kinetics with respect to concentration;

$$S_2 = \int_0^t a\ C\ dt \qquad (10)$$

where a is the instantaneous rate coefficient (hr^{-1})

(4) the soil column behaves as if it were infinitely long with constant water velocity; and

(5) the total sorption is the sum of the reversibly sorbed and irreversibly precipitated following the equation

$$S = S_1 + S_2.$$

One solution has been proposed as

$$\frac{C}{C_T} = \frac{1}{2} \left(\text{erf} \frac{X + X_o - Vt/R}{2\sqrt{D/R}} - \text{erf} \frac{X - Vt/R}{2\sqrt{D/R}} \right) \qquad (11)$$

where:

erf = error function

$C_T = C_o \exp(-\rho ta/\theta R)$

C_o = initial concentration (mg/l)

R = retardation factor = $1 + \frac{\rho}{\theta} k_d$

X_o = depth of an equivalent length of P describing the applied chemical $X_o = t_o V/R$

t_o = time P applied at surface (hrs)

The second approach is a numerical solution to Equation 9 developed with appropriate boundary conditions by Mansell et al. (1977) where:

$$\frac{\partial S}{\partial t} = K_r (K C^N - S) \qquad (12)$$

The third model considered follows the kinetic Equation 8 which was solved using the method of characteristics (Enfield et al. 1981).

RESULTS AND DISCUSSION

The models were tested using data already in the literature. Model coefficients were developed where necessary when they were different than those presented in the literature. Figure 4 is the sorption data presented by Mansell et al. (1977) for Oldsmar fine sand (alfic arenic haplaquod) A1 horizon. Figure 5 shows an adaptation of the kinetic data from the same soil sample. Ignoring the initial rapid sorption, the rate of reaction appears to be approximately 0.02 hr^{-1}. Figure 6 shows experimental data of Mansell et al. for a 10 cm column of soil. The experimental conditions are described in Tables 2 and 3. The projection in Figure 6 are assuming a Freundlich sorption which has not come to equilibrium after the development of Mansell. The projections in Figure 7 are for the model projections assuming Langmuir sorption and first-order precipitation where it is assumed the concentration of P at equilibrium will be 0.01 mg/l and dispersion negligible. Two different precipitation rate coefficients are used. When the projection is based on the rate observed in the batch sorption study, one projects the concentration should have been greater than observed. The fast rate (0.1 hr^{-1}) is a "calibrated" value to make the projection look as close as possible to experimental data. Figure 8 shows the same experimental data but assumes linear sorption and first-order dissipation or irreversible precipitation. The linear partition coefficient k_d was "calibrated" to the column data and not related to any sorption study. Two curves are projected in Figure 8: one is for the case with no dispersion and the other is for dispersion as reported by Mansell. The rate of irreversible precipitation is 0.1 as was used in Figure 7. Figure 9 is similar to the two previous, but the flow rate is much lower with interstitial velocities of 1.2 cm/hr versus

9.27 cm/hr. The results obtained are comparable. At the lower flow rate, the rate of precipitation calculated from the batch sorption studies appears to describe the column breakthrough data much more precisely than was observed in the previous sample. The models appear to give comparable results. There are two significant differences that should be delineated.

(1) When irreversible precipitation is assumed, the solution concentration returns to zero shortly after the pulse of P passes, while with reversible precipitation the concentration is not zero until all the P precipitated is returned to solution.

(2) The projected time for breakthrough was adequately described using batch studies for the Freundlich or Langmuir equation but the linear sorption coefficient was "calibrated" to the column data. This simplification allows using a less operationally complex model but judgment is needed to understand the consequences of this simplification.

A limitation of the model presented based on Langmuir sorption is its lack of considering the effect of dispersion. This is not a significant problem if one is considering the response of a large more or less continuous addition of P as with waste water applications to the land. However, when one is applying an application of P similar to that in normal fertility practice, where the concentration is high for a very short period of time, the influence of dispersion on the concentration of P in the soil solution is dramatic (Enfield and Carsel, 1980). If the activity in the soil solution controls the plant availability rather than the amount in the solid phase, the Langmuir model would need to include dispersion to be effective for most agricultural situations.

The term in either model attempting to describe precipitation which appears to be most significant in estimating solution concentration is the rate coefficient. There have been attempts to relate this coefficient to other readily measurable parameters. Figure 10 shows a collection of these results as a function of pH. Enfield et al. (1981) showed pH vs. rate coefficient to be statistically significant at the 0.005 level in calcareous soils. Neutral soils appear to have the slowest rate of reaction as well as the greatest solubility for the phosphate chemicals.

CONCLUSIONS

Several different types of models have been proposed to describe the disappearance of P from soil solutions. None of the models have claimed to fully explain the actual mechanisms for P reaction with soil. Of the kinetic models that have been proposed, two general approaches appear to approximate the majority of reported data: those where sorption sites are assumed to not be at equilibrium having different rates of sorption and desorption; and those based on kinetics of precipitation and dissolution. Although an understanding of the mechanisms involved is incomplete, approximations can be made for the mobility and activity of P in soil solution.

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LIST OF FIGURES

Figure 1. Stability diagram for selected phosphate compounds. Fe-P is the stability line for strengite ($FePO_4 \cdot 2H_2O$) in equilibrium with $Fe(OH)_3$. Al-P is the stability line for variscite ($AlPO_4 \cdot 2H_2O$) in equilibrium with amorphous $Al(OH)_3$. The calcium phosphates dicalcium phosphate dyhydrate (DCPD) $Ca(H_2PO_4)_2 \cdot H_2O$, octacalcium phosphate (OCP) $Ca_8H_2(PO_4)_6 \cdot 5H_2O$, tricalcium phosphate (TCP) $Ca_3(PO_4)_2$, and hydroxyapatite (HA) $Ca_{10}(OH)_2\ (PO_4)_6$ are calculated for 0.01 M Ca and an ionic strength of 0.026 mole $liter^{-1}$.

Figure 2. Successive sorption (desorption of phosphorus by a single soil sample (personal communication Dr. Arne Stuanes, Norwegian Forest Research Institute, Aas, NLH 1981).

Figure 3. Timed response for sorption of phosphorus by a single soil sample.

Figure 4. Sorption isotherm for Oldsmar fine sand after 24 of equilibration adapted from Mansell et al. 1977.

Figure 5. Kinetics of sorption for Oldsmar fine sand with an initial solution concentration of 100 mg/l adapted from Mansell et al. 1977.

Figure 6. Measured and projected breakthrough curves assuming Freundlich sorption which had not reached equilibrium adapted from Mansell et al. 1977.

Figure 7. Measured and projected breakthrough curve assuming instantaneous Langmuir sorption and chemical precipitation and disolution.

Figure 8. Measured and projected breakthrough curves assuming instantaneous linear sorption and irreversible precipitation.

Figure 9. Measured and projected breakthrough curves assuming either Langmuir or linear sorption plus precipitation.

Figure 10. Specific rate as a function of pH. Data for Kaolinite and Al_2O_3 is referred to the liquid phase; data on acid soils and hydroxyapatite is referred to the solid phase.

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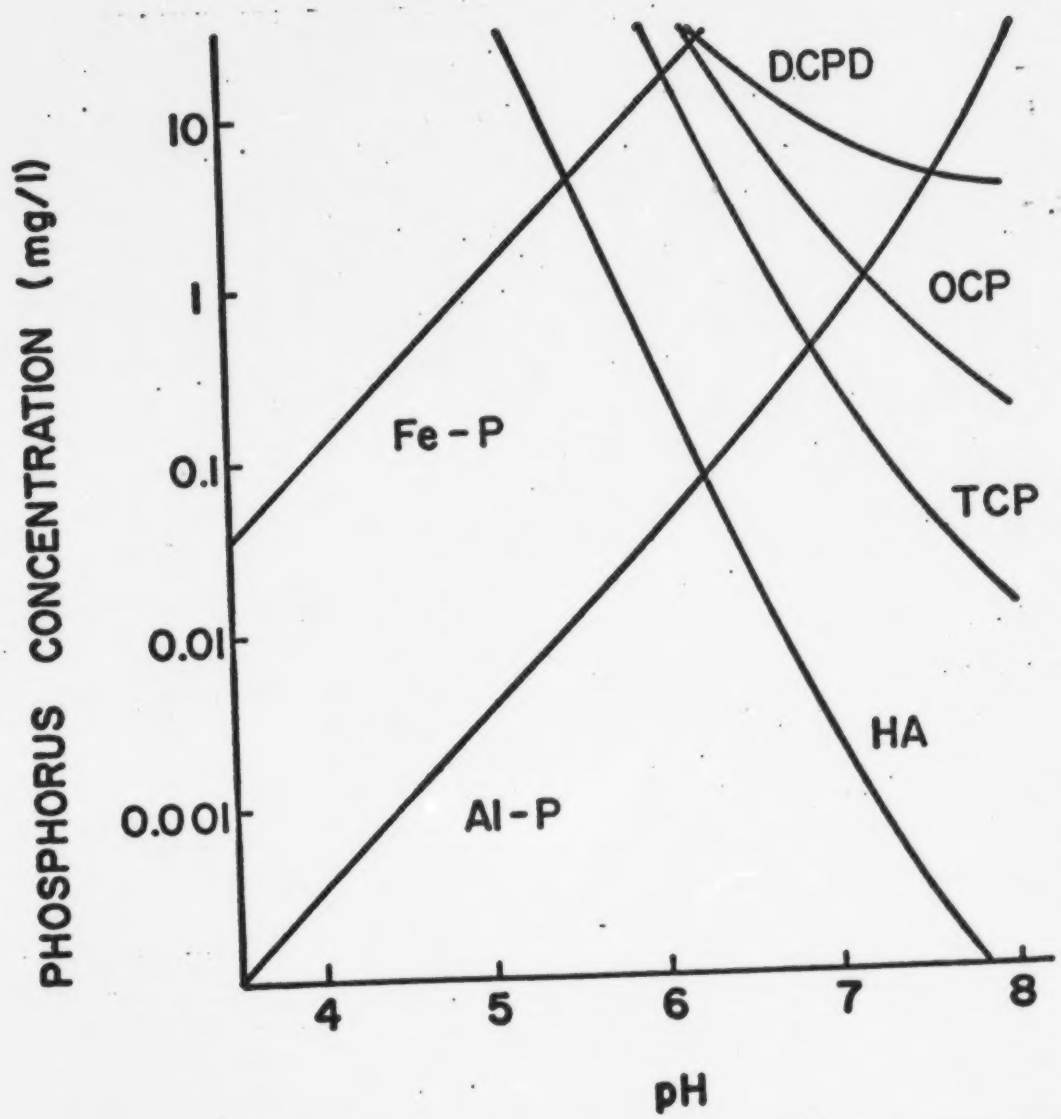
Table 1. Reported Gibbs Free Energy of Selected Compounds.

Table 2. Physical Parameters for the Soil Columns.

Table 3. Chemical Parameters for the Soil Columns.


PHOSPHORUS CONCENTRATION (mg/l)
10
1
0.1
0.01
0.001
DCPD
OCP
TCP
HA
Fe-P
Al-P
4
5
6
7
8
pH

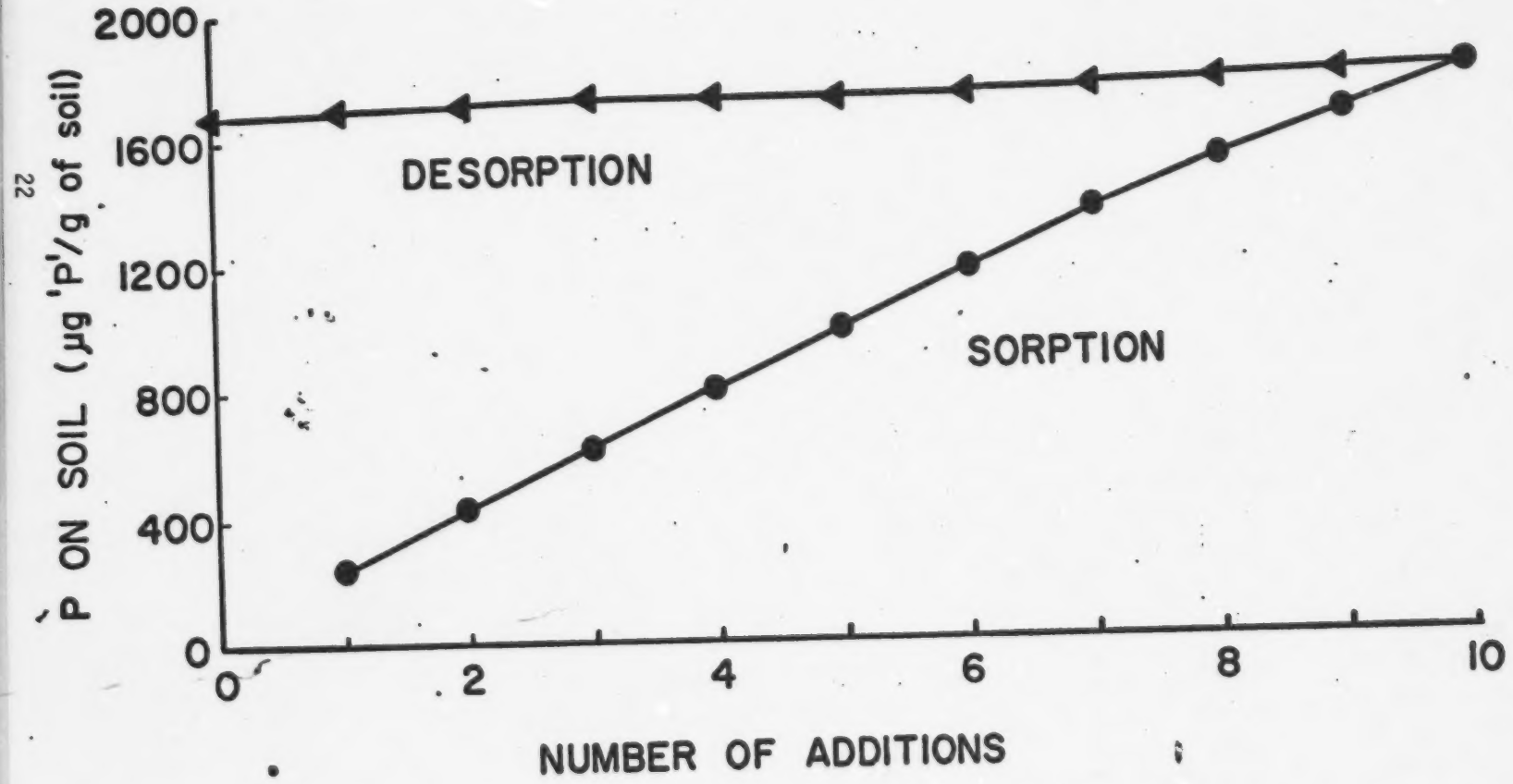
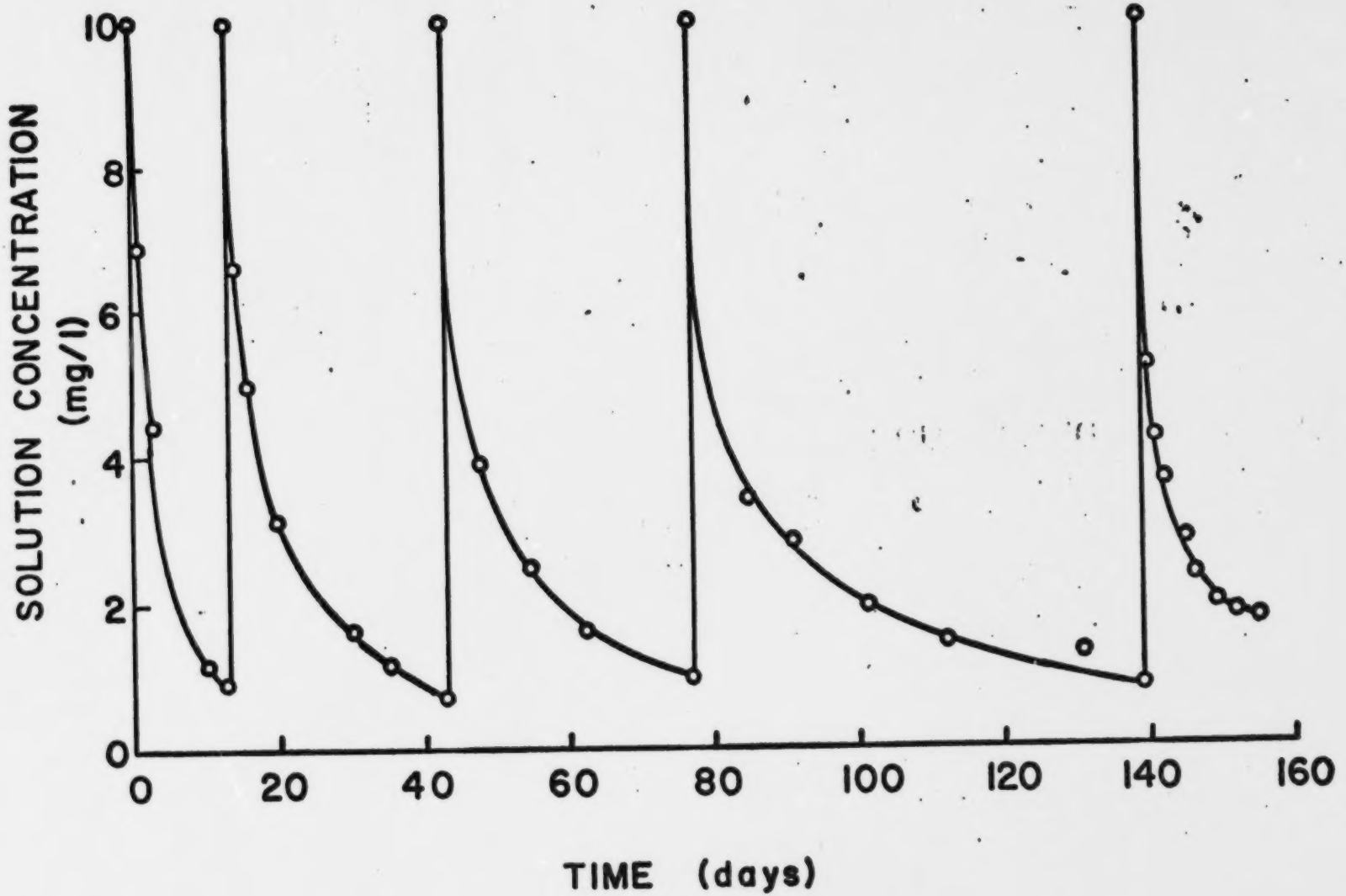

2000
1600
1200
800
400
0
P ON SOIL (µg 'P'/g of soil)
DESORPTION
SORPTION
0
2
4
6
8
10
NUMBER OF ADDITIONS

FIG 2


SOLUTION CONCENTRATION (mg/l)
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2
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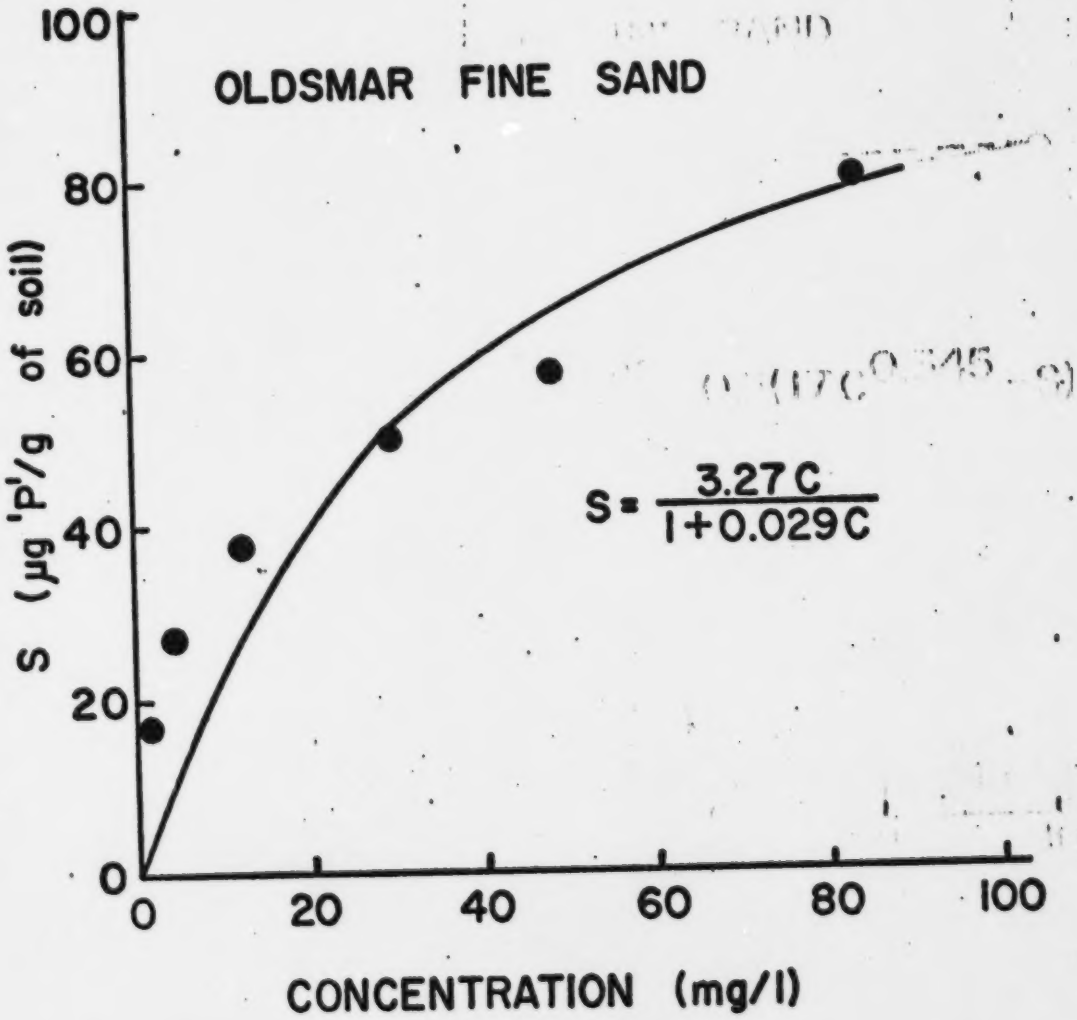
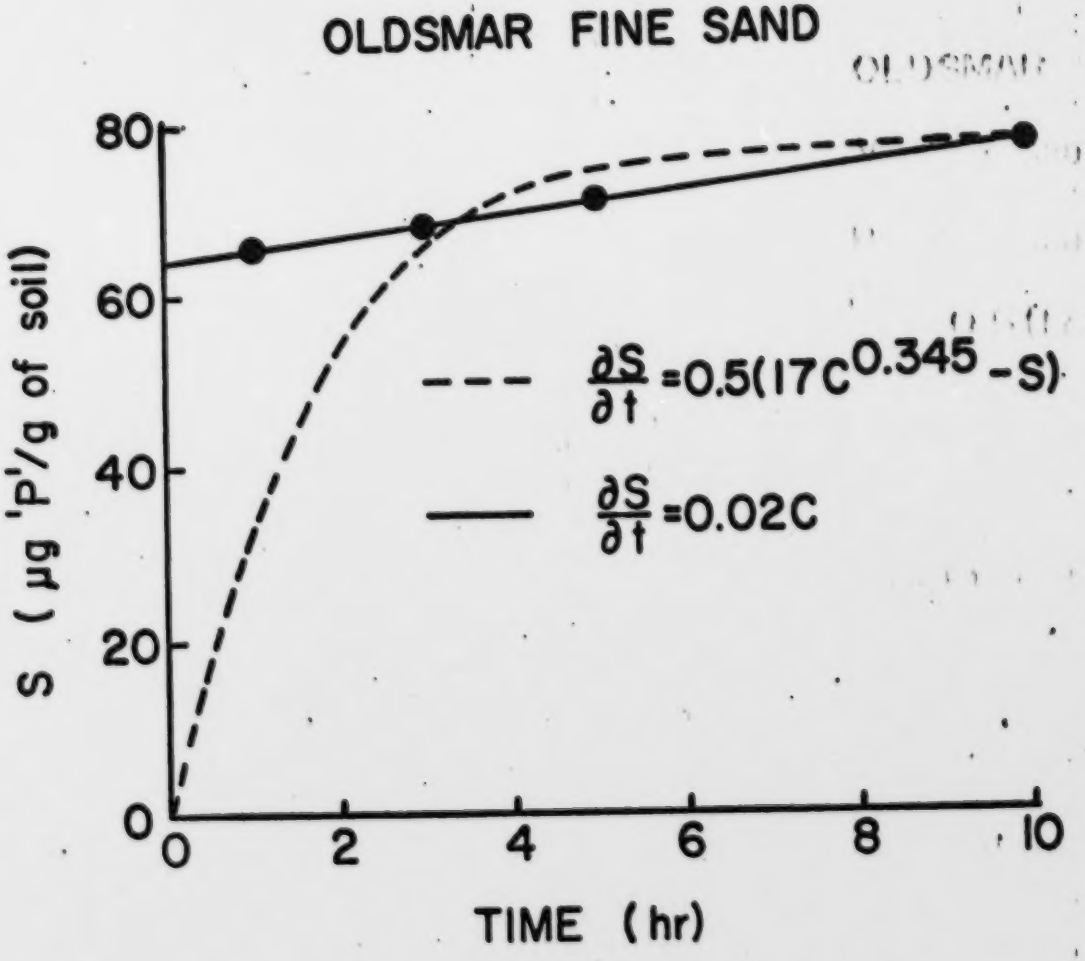
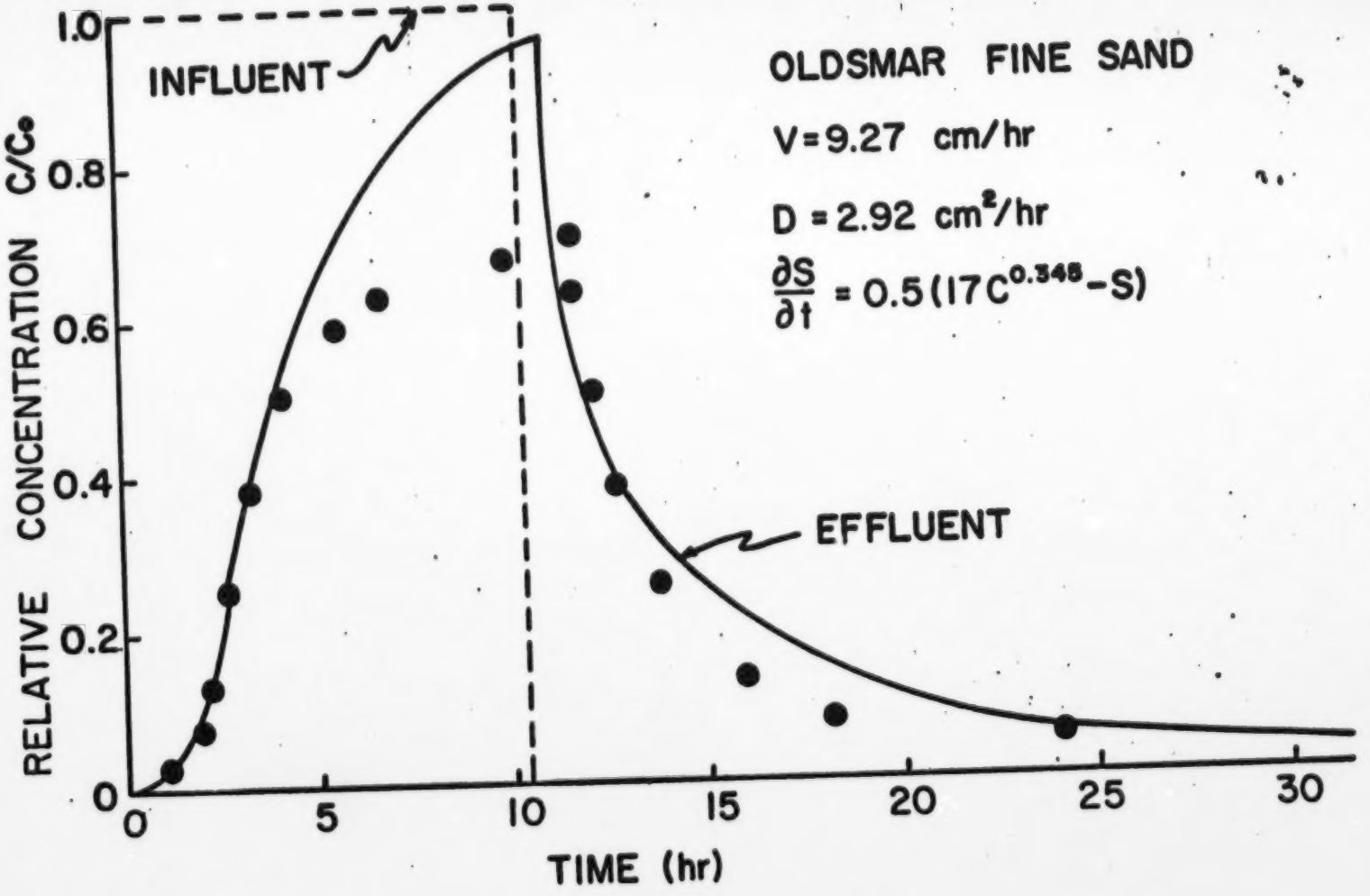
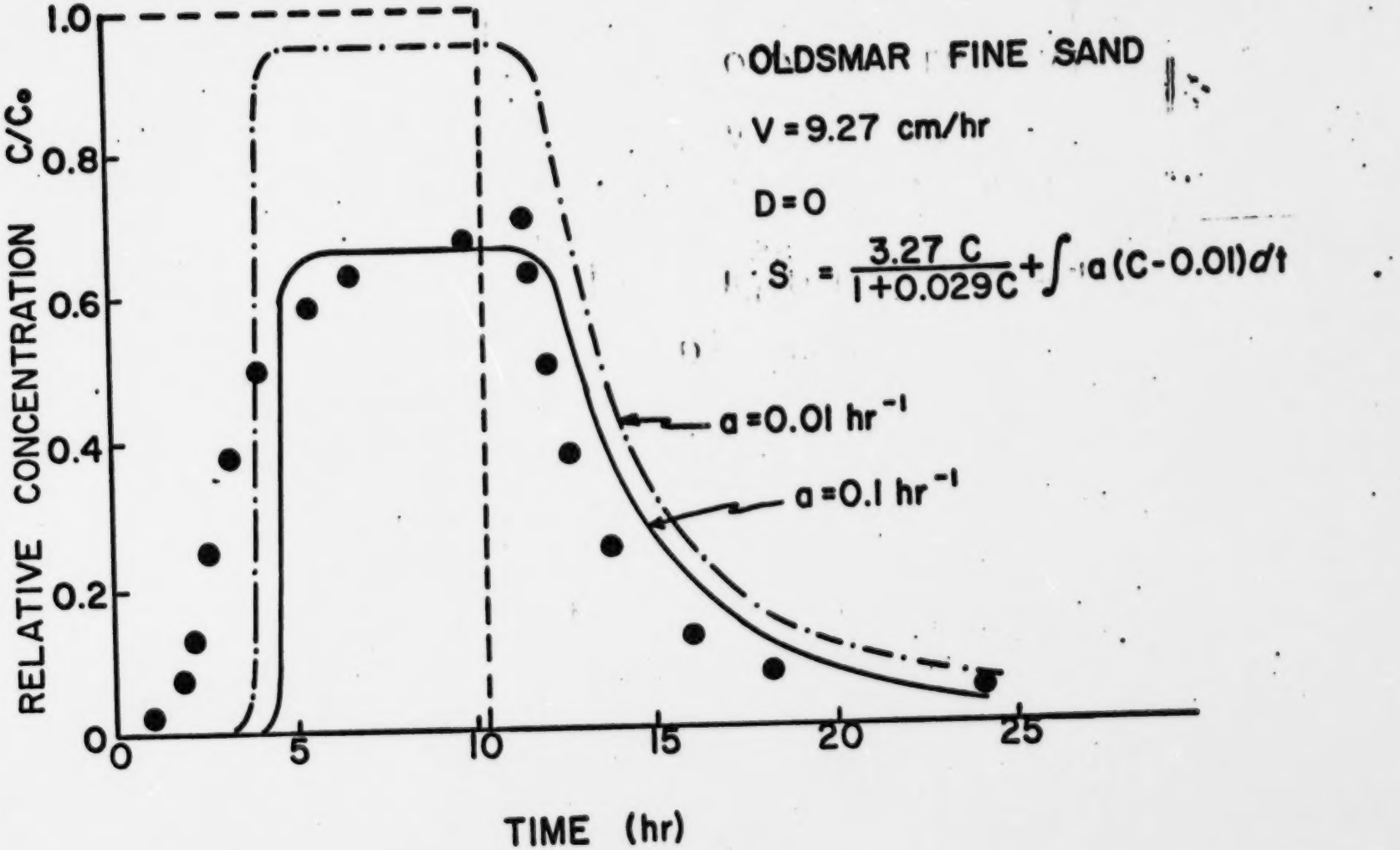
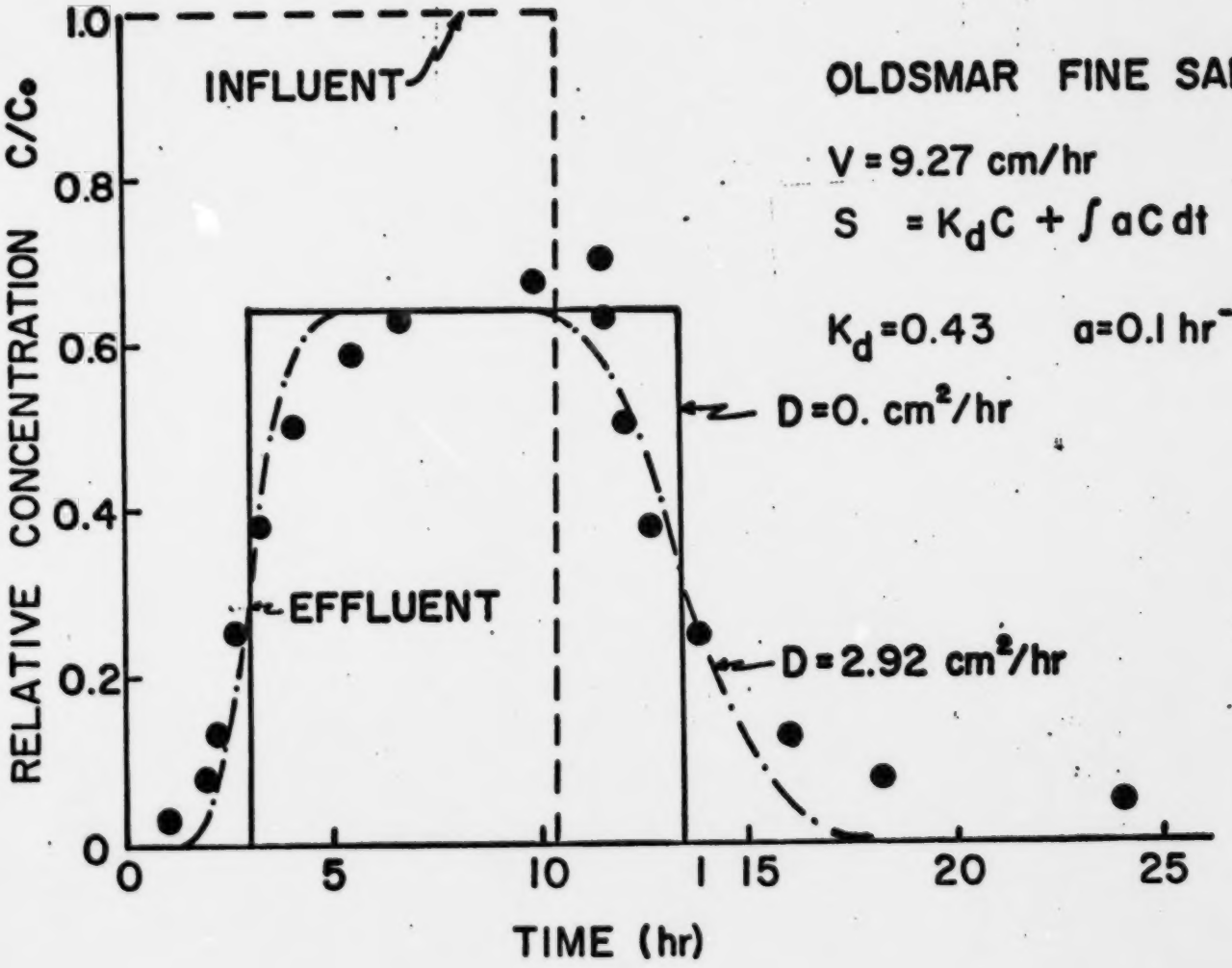
6
8
10
0
20
40
60
80
100
120
140
160
TIME (days)

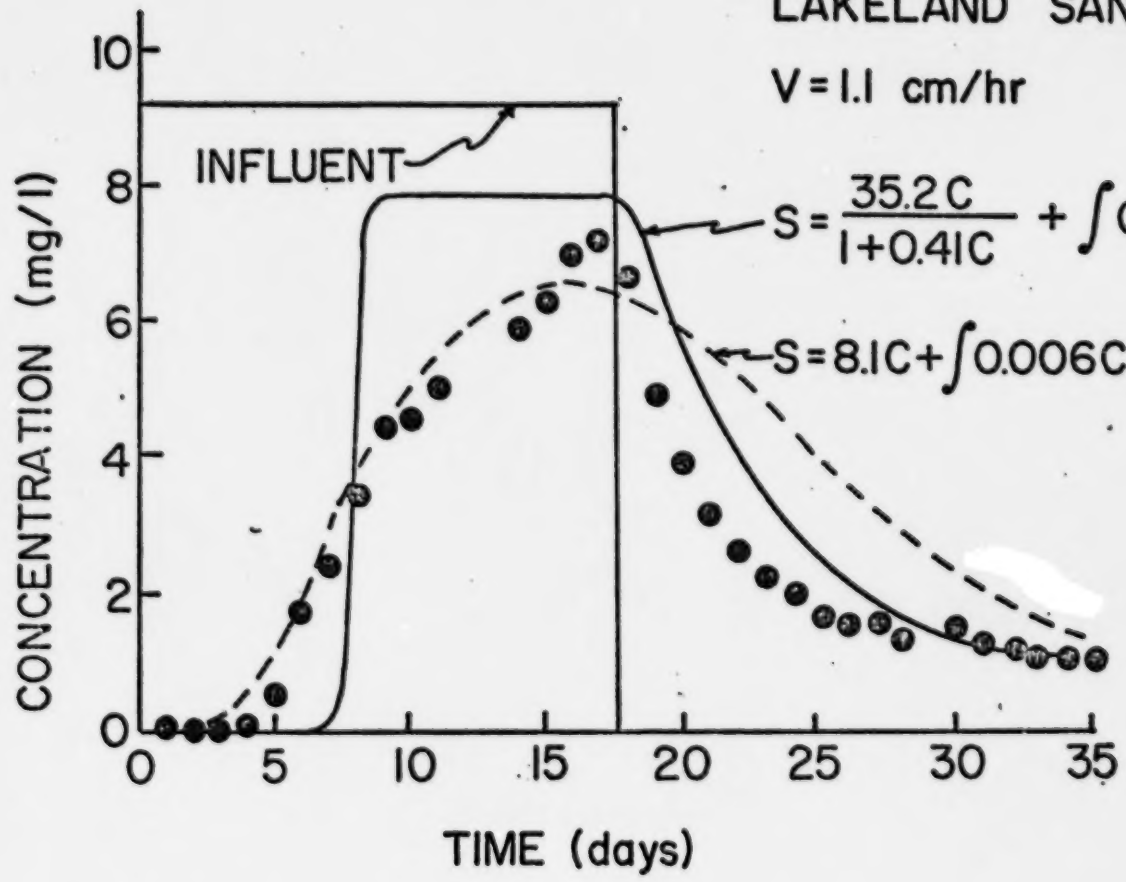

OLDSMAR FINE SAND
S (µg 'P'/g of soil)
CONCENTRATION (mg/l)
S = 3.27C / 1+0.029C
0
20
40
60
80
100


OLDSMAR FINE SAND
S (µg 'P'/g of soil)
TIME (hr)
0
20
40
60
80
0
2
4
6
8
10
$\frac{\partial S}{\partial t}=0.5(17C^{0.345}-S)$
$\frac{\partial S}{\partial t}=0.02C$


INFLUENT
EFFLUENT
OLDSMAR FINE SAND
V = 9.27 cm/hr
D = 2.92 cm²/hr
∂S/∂t = 0.5(17C^0.348 - S)
RELATIVE CONCENTRATION C/C₀
TIME (hr)
1.0
0.8
0.6
0.4
0.2
0
0
5
10
15
20
25
30


OLDSMAR FINE SAND
V = 9.27 cm/hr
D = 0
$S = \frac{3.27\,C}{1+0.029C} + \int a(C-0.01)dt$
a = 0.01 hr⁻¹
a = 0.1 hr⁻¹
RELATIVE CONCENTRATION C/C₀
TIME (hr)
0
0.2
0.4
0.6
0.8
1.0
5
10
15
20
25


OLDSMAR FINE SAND
V = 9.27 cm/hr
S = $K_d C + \int \alpha C\, dt$
K_d = 0.43 α = 0.1 hr^{-1}
D = 0. cm^2/hr
D = 2.92 cm^2/hr
INFLUENT
EFFLUENT
RELATIVE CONCENTRATION C/C₀
TIME (hr)
1.0
0.8
0.6
0.4
0.2
0
0
5
10
15
20
25


LAKELAND SAND
V = 1.1 cm/hr
INFLUENT
$S = \frac{35.2C}{1+0.41C} + \int 0.006(C-0.01)dt$ D=0
$S = 8.1C + \int 0.006C dt$ D=0.8
CONCENTRATION (mg/l)
TIME (days)
0
2
4
6
8
10
0
5
10
15
20
25
30
35

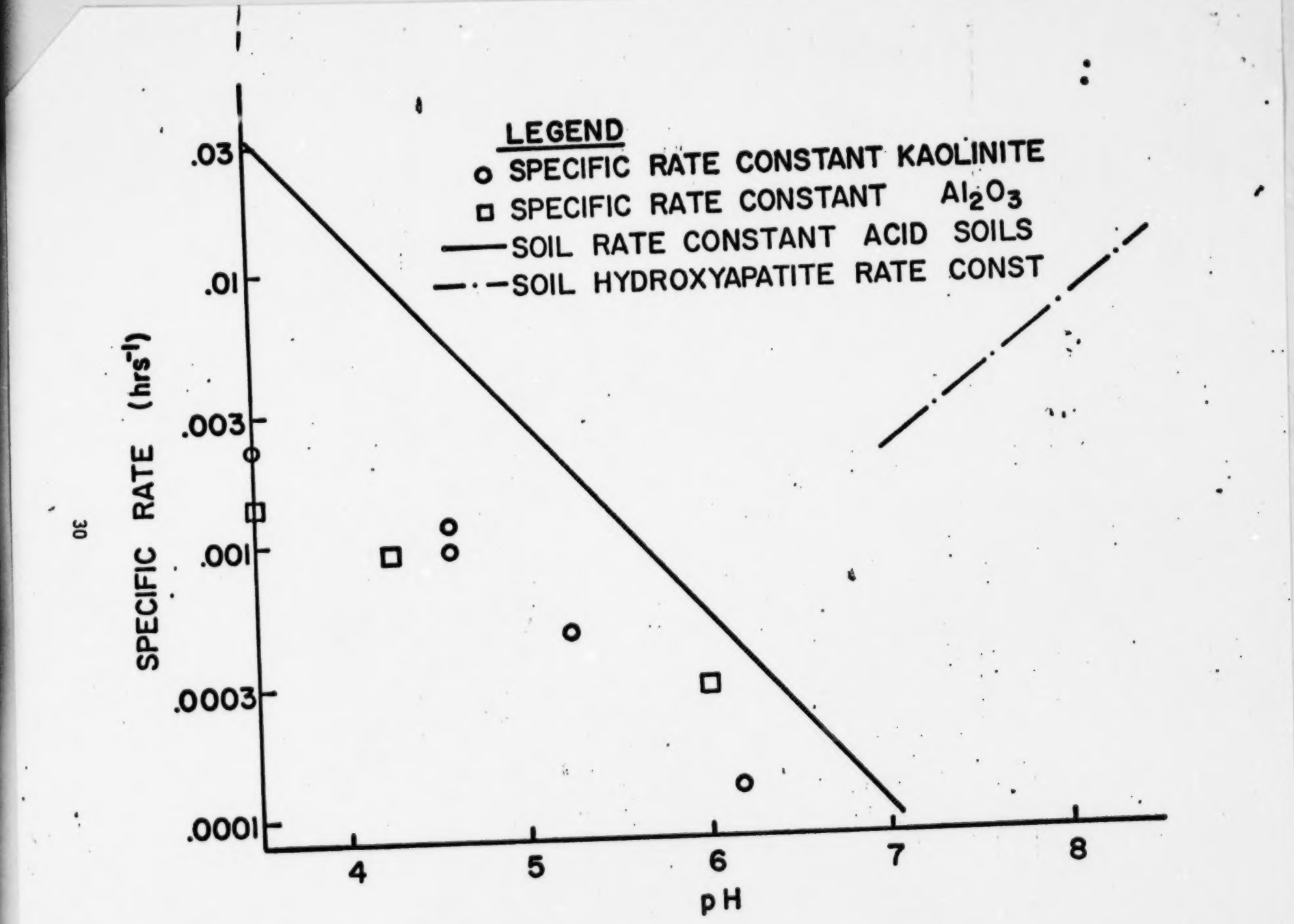

LEGEND
SPECIFIC RATE CONSTANT KAOLINITE
SPECIFIC RATE CONSTANT Al_2O_3
SOIL RATE CONSTANT ACID SOILS
SOIL HYDROXYAPATITE RATE CONST
SPECIFIC RATE (hrs^{-1})
.03
.01
.003
.001
.0003
.0001
4
5
6
7
8
pH

TABLE 1. REPORTED GIBBS FREE ENERGY OF SELECTED COMPOUNDS*

	ΔG_F°
Al^{3+}	-117.33
$AlOH^{2+}$	-167.17
$Al(OH)_2^+$	-218.02
$Al(OH)_3^\circ$	-266.94
$Al(OH)_3$ (amorphous)	-274.21
α - $Al(OH)_3$ (bayerite)	-275.78
γ - $Al(OH)_3$ (gibbsite)	-276.43
$Al(OH)_3$ (nordstrandite)	-276.3
$AlPO_4$ (berlinite)	-388.50
$AlPO_4 \cdot 2H_2O$ (variscite)	-505.97
Ca^{2+}	-132.52
$CaCO_3^\circ$	-263.00
$CaCO_3$ (calcite)	-270.18
$CaPO_4^-$	-386.51
$CaHPO_4^\circ$	-398.29
$CaH_2PO_4^+$	-406.28
$CaHPO_4$ (monetite)	-403.96
$CaHPO_4 \cdot 2H_2O$ (brushite)	-516.89
$Ca(H_2PO_4)_2 \cdot H_2O$ (c)	-734.48
α - $Ca_3(PO_4)_2$ (c)	-922.70
β - $Ca_3(PO_4)_2$ (tricalcium phosphate)	-927.37
$Ca_8H_2(PO_4)_6 \cdot 5H_2O$ (octacalcium phosphate)	-2942.62
$Ca_{10}(OH)_2(PO_4)_6$ (hydroxyapatite)	-3030.24
$FeOH^+$	-69.29
$FeOH^{2+}$	-57.72

(continued)

TABLE 1. (continued)

	ΔG_F^o
$Fe(OH)_2$ (c)	-117.58
$Fe(OH)_3$ (amorphous)	-169.25
$Fe(OH)_3$ (soil)	-170.40
α - FeOOH (geothite)	-117.42
$FePO_4 \cdot 2H_2O$ (strengite)	-398.59
$Fe_3(PO_4)_2 \cdot 8H_2O$ (vivianite)	-1058.36
PO_4^{3-}	-245.18
HPO_4^{2-}	-262.03
$H_2PO_4^-$	-271.85
$H_3PO_4^o$	-274.78
H_2O (l)	-56.69

* From Sadiq and Lindsay (20)

TABLE 2. PHYSICAL PARAMETERS FOR THE SOIL COLUMNS

Soil	Interstitial velocity cm/hr	Water Content (θ)(cc/cc)	Bulk Density ρ (g/cm^3)	Application Rate cm/hr	Time of "slug" hrs	Dispersion Coefficient cm^2/hr	Initial Concen- mg/l	Column length cm
Oldsmar †	9.27	0.34	1.39	3.15	10.4	2.92	100	10
Lakeland ††	1.0	0.26	1.5	0.3	408.	0.8	9.2	5

† Data adopted from Mansell et al. 1977

†† Data adopted from Enfield et al. 1981

TABLE 3. CHEMICAL PARAMETERS FOR THE SOIL COLUMNS

Soil	K_d*	Langmuir Max µg/g	Langmuir b	Rate hr^{-1}	Freundlich n	Freundlich K
Oldsmar	0.43	112.9§	0.029§	0.022§	0.345	17
Lakeland	8.1	86†	0.41†	0.006†		

* Calibrated value based on correct breakthrough time not related to batch sorption studies

§ Calculated from 7 day adsorption data and kinetic data reported by Mansell et al. 1977

† Reported by Enfield et al. 1981